

conditions, *no molecules exist in a sample of an ionic compound*. A piece of sodium chloride, for example, is a continuous array of oppositely charged sodium and chloride ions, *not* a collection of individual “sodium chloride molecules.”

Another key distinction exists between the particles attracting each other. Covalent bonding involves the mutual attraction between two (positively charged) nuclei and the two (negatively charged) electrons that reside between them. Ionic bonding involves the mutual attraction between positive and negative ions.

Polyatomic Ions: Covalent Bonds Within Ions Many ionic compounds contain **polyatomic ions**, which consist of two or more atoms bonded *covalently* and have a net positive or negative charge. For example, the ionic compound calcium carbonate is an array of polyatomic carbonate anions and monatomic calcium cations attracted to each other. The carbonate ion consists of a carbon atom covalently bonded to three oxygen atoms, and two additional electrons give the ion its 2[−] charge (Figure 2.15). In many reactions, a polyatomic ion stays together as a unit.

SECTION SUMMARY

Although a few elements occur uncombined in nature, the great majority exist in compounds. Ionic compounds form when a metal *transfers electrons* to a nonmetal, and the resulting positive and negative ions attract each other to form a three-dimensional array. In many cases, metal atoms lose and nonmetal atoms gain enough electrons to attain the same number of electrons as in atoms of the nearest noble gas. Covalent compounds form when elements, usually nonmetals, *share electrons*. Each covalent bond is an electron pair mutually attracted by two atomic nuclei. Monatomic ions are derived from single atoms. Polyatomic ions consist of two or more covalently bonded atoms that have a net positive or negative charge due to a deficit or excess of electrons.

2.8 COMPOUNDS: FORMULAS, NAMES, AND MASSES

Names and formulas of compounds form the vocabulary of the chemical language. In this discussion, you'll learn the names and formulas of ionic and simple covalent compounds and how to calculate the mass of a unit of a compound from its formula.

Types of Chemical Formulas

In a **chemical formula**, element symbols and numerical subscripts show the type and number of each atom present in the smallest unit of the substance. There are several types of chemical formulas for a compound:

1. The **empirical formula** shows the *relative* number of atoms of each element in the compound. It is the simplest type of formula and is derived from the masses of the component elements. For example, in hydrogen peroxide, there is 1 part by mass of hydrogen for every 16 parts by mass of oxygen. Therefore, the empirical formula of hydrogen peroxide is HO: one H atom for every O atom.
2. The **molecular formula** shows the *actual* number of atoms of each element in a molecule of the compound. The molecular formula of hydrogen peroxide is H₂O₂; there are two H atoms and two O atoms in each molecule.
3. A **structural formula** shows the number of atoms and *the bonds between them*; that is, the relative placement and connections of atoms in the molecule. The structural formula of hydrogen peroxide is H—O—O—H; each H is bonded to an O, and the O's are bonded to each other.

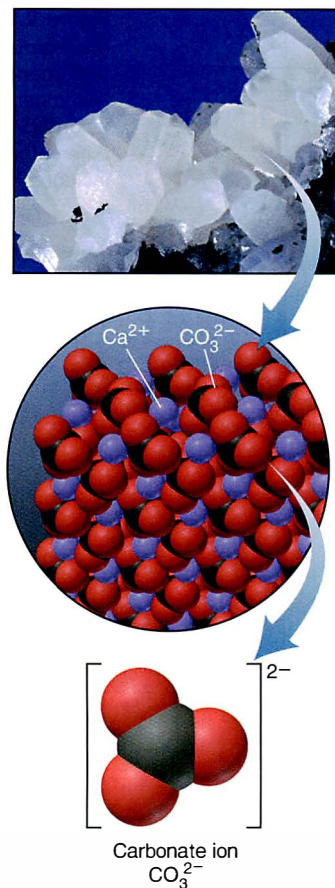


Figure 2.15 A polyatomic ion. Calcium carbonate is a three-dimensional array of monatomic calcium cations (*purple spheres*) and polyatomic carbonate anions. As the bottom structure shows, each carbonate ion consists of four covalently bonded atoms.

Figure 2.16 Some common monatomic ions of the elements. Main-group elements usually form a single monatomic ion. Note that members of a group have ions with the same charge. [Hydrogen is shown as both the cation H^+ in Group 1A(1) and the anion H^- in Group 7A(17).] Many transition elements form two different monatomic ions. (Although Hg_2^{2+} is a diatomic ion, it is included for comparison with Hg^{2+} .)

	1A (1)	2A (2)																7A (17)	8A (18)
1	H^+																	H^-	
2	Li^+																	F^-	
3	Na^+	Mg^{2+}	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	9B (9)	10B (10)	1B (11)	2B (12)	Al^{3+}				S^{2-}	Cl^-	
4	K^+	Ca^{2+}				Cr^{2+} Cr^{3+}	Mn^{2+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}		Cu^+ Cu^{2+}	Zn^{2+}						Br^-	
5	Rb^+	Sr^{2+}									Ag^+	Cd^{2+}		Sn^{2+} Sn^{4+}				I^-	
6	Cs^+	Ba^{2+}										Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}					
7																			

Names and Formulas of Ionic Compounds

All ionic compound names give the positive ion (cation) first and the negative ion (anion) second. Here are some points to note about ion charges:

- Members of a periodic table group have the same ionic charge; for example, Li, Na, and K are all in Group 1A and all have a 1+ charge.
- For A-group cations, ion charge = group number: for example, Na^+ is in Group 1A, Ba^{2+} in Group 2A. (Exceptions in Figure 2.16 are Sn^{2+} and Pb^{2+} .)
- For anions, ion charge = group number minus 8: for example, S is in Group 6A ($6 - 8 = -2$), so the ion is S^{2-} . Table 2.3 lists some of the more common monatomic ions.

Compounds Formed from Monatomic Ions Let's first consider how to name binary ionic compounds, those composed of ions of two elements.

- The name of the cation is the same as the name of the metal. Many metal names end in *-ium*.
- The name of the anion takes the root of the nonmetal name and adds the suffix *-ide*.

For example, the anion formed from bromine is named bromide (brom+ide). Therefore, the compound formed from the metal calcium and the nonmetal bromine is named *calcium bromide*.

SAMPLE PROBLEM 2.5 Naming Binary Ionic Compounds

Problem Name the ionic compound formed from the following pairs of elements:

- (a) Magnesium and nitrogen (b) Iodine and cadmium
(c) Strontium and fluorine (d) Sulfur and cesium

Plan The key to naming a binary ionic compound is to recognize which element is the metal and which is the nonmetal. When in doubt, check the periodic table. We place the cation name first, add the suffix *-ide* to the nonmetal root, and place the anion name last.

Solution (a) Magnesium is the metal; *nitr-* is the nonmetal root: magnesium nitride

(b) Cadmium is the metal; *iod-* is the nonmetal root: cadmium iodide

Table 2.3 Common Monatomic Ions*

Charge	Formula	Name
Cations		
1+	H^+	hydrogen
	Li^+	lithium
	Na^+	sodium
	K^+	potassium
	Cs^+	cesium
	Ag^+	silver
2+	Mg^{2+}	magnesium
	Ca^{2+}	calcium
	Sr^{2+}	strontium
	Ba^{2+}	barium
	Zn^{2+}	zinc
	Cd^{2+}	cadmium
3+	Al^{3+}	aluminum
Anions		
1-	H^-	hydride
	F^-	fluoride
	Cl^-	chloride
	Br^-	bromide
2-	I^-	iodide
	O^{2-}	oxide
	S^{2-}	sulfide
3-	N^{3-}	nitride

*Listed by charge; those in **boldface** are most common.

(c) Strontium is the metal; *fluor-* is the nonmetal root: **strontium fluoride** (Note the spelling is *fluoride*, not *flouride*.)

(d) Cesium is the metal; *sulf-* is the nonmetal root: **cesium sulfide**

FOLLOW-UP PROBLEM 2.5 For the following ionic compounds, give the name and periodic table group number of each of the elements present: (a) zinc oxide; (b) silver bromide; (c) lithium chloride; (d) aluminum sulfide.

Ionic compounds are arrays of oppositely charged ions rather than separate molecular units. Therefore, we write a formula for the **formula unit**, which gives the *relative* numbers of cations and anions in the compound. Thus, ionic compounds generally have only empirical formulas.* The compound has zero net charge, so the positive charges of the cations must balance the negative charges of the anions. For example, calcium bromide is composed of Ca^{2+} ions and Br^- ions; therefore, two Br^- balance each Ca^{2+} . The formula is CaBr_2 , not Ca_2Br . In this and all other formulas,

- The subscript refers to the element *preceding* it.
- The *subscript 1 is understood* from the presence of the element symbol alone (that is, we do not write Ca_1Br_2).
- The charge (without the sign) of one ion becomes the subscript of the other:



Reduce the subscripts to the smallest whole numbers that retain the ratio of ions. Thus, for example, from the ions Ca^{2+} and O^{2-} we have Ca_2O_2 , which we reduce to the formula CaO (but see the footnote).

SAMPLE PROBLEM 2.6 Determining Formulas of Binary Ionic Compounds

Problem Write empirical formulas for the compounds named in Sample Problem 2.5.

Plan We write the empirical formula by finding the smallest number of each ion that gives the neutral compound. These numbers appear as *right subscripts* to the element symbol.

Solution

(a) Mg^{2+} and N^{3-} ; three Mg^{2+} ions (6+) balance two N^{3-} ions (6-): **Mg_3N_2**

(b) Cd^{2+} and I^- ; one Cd^{2+} ion (2+) balances two I^- ions (2-): **CdI_2**

(c) Sr^{2+} and F^- ; one Sr^{2+} ion (2+) balances two F^- ions (2-): **SrF_2**

(d) Cs^+ and S^{2-} ; two Cs^+ ions (2+) balance one S^{2-} ion (2-): **Cs_2S**

Comment Note that ion charges do *not* appear in the compound formula. That is, for cadmium iodide, we do *not* write $\text{Cd}^{2+}\text{I}_2^-$.

FOLLOW-UP PROBLEM 2.6 Write the formulas of the compounds named in Follow-up Problem 2.5.

Compounds with Metals That Can Form More Than One Ion Many metals, particularly the transition elements (B groups), can form more than one ion, each with its own particular charge. Table 2.4 (on the next page) lists some examples, and Figure 2.16 shows their placement in the periodic table. Names of compounds containing these elements include a *Roman numeral within parentheses* immediately after the metal ion's name to indicate its ionic charge. For example, iron can form Fe^{2+} and Fe^{3+} ions. The two compounds that iron forms with

*Compounds of the mercury(I) ion, such as Hg_2Cl_2 , and peroxides of the alkali metals, such as Na_2O_2 , are the only two common exceptions. Their empirical formulas are HgCl and NaO , respectively.

Table 2.4 Some Metals That Form More Than One Monatomic Ion*

Element	Ion Formula	Systematic Name	Common (Trivial) Name
Chromium	Cr^{2+}	chromium(II)	chromous
	Cr^{3+}	chromium(III)	chromic
Cobalt	Co^{2+}	cobalt(II)	
	Co^{3+}	cobalt(III)	
Copper	Cu^{+}	copper(I)	cuprous
	Cu^{2+}	copper(II)	cupric
Iron	Fe^{2+}	iron(II)	ferrous
	Fe^{3+}	iron(III)	ferric
Lead	Pb^{2+}	lead(II)	
	Pb^{4+}	lead(IV)	
Mercury	Hg_2^{2+}	mercury(I)	mercurous
	Hg^{2+}	mercury(II)	mercuric
Tin	Sn^{2+}	tin(II)	stannous
	Sn^{4+}	tin(IV)	stannic

*Listed alphabetically by metal name; those in **boldface** are most common.

Table 2.5 Common Polyatomic Ions*

Formula	Name
Cations	
NH_4^+	ammonium
H_3O^+	hydronium
Anions	
CH_3COO^- (or $\text{C}_2\text{H}_3\text{O}_2^-$)	acetate
CN^-	cyanide
OH^-	hydroxide
ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate
NO_2^-	nitrite
NO_3^-	nitrate
MnO_4^-	permanganate
CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate (or bicarbonate)
CrO_4^{2-}	chromate
$\text{Cr}_2\text{O}_7^{2-}$	dichromate
O_2^{2-}	peroxide
PO_4^{3-}	phosphate
HPO_4^{2-}	hydrogen phosphate
H_2PO_4^-	dihydrogen phosphate
SO_3^{2-}	sulfite
SO_4^{2-}	sulfate
HSO_4^-	hydrogen sulfate (or bisulfate)

***Boldface** ions are most common.

chlorine are FeCl_2 , named iron(II) chloride (spoken “iron two chloride”), and FeCl_3 , named iron(III) chloride.

In common names, the Latin root of the metal is followed by either of two suffixes:

- The suffix **-ous** for the ion with the lower charge
- The suffix **-ic** for the ion with the higher charge

Thus, iron(II) chloride is also called ferr**ous** chloride and iron(III) chloride is ferr**ic** chloride. (You can easily remember this naming relationship because there is an *o* in **-ous** and *lower*, and an *i* in **-ic** and *higher*.)

SAMPLE PROBLEM 2.7 Determining Names and Formulas of Ionic Compounds of Elements That Form More Than One Ion

Problem Give the systematic names for the formulas or the formulas for the names of the following compounds: (a) tin(II) fluoride; (b) CrI_3 ; (c) ferric oxide; (d) CoS .

Solution (a) Tin(II) is Sn^{2+} ; fluoride is F^- . Two F^- ions balance one Sn^{2+} ion: tin(II) fluoride is SnF_2 . (The common name is stannous fluoride.)

(b) The anion is I^- , iodide, and the formula shows three I^- . Therefore, the cation must be Cr^{3+} , chromium(III): CrI_3 is chromium(III) iodide. (The common name is chromic iodide.)

(c) Ferric is the common name for iron(III), Fe^{3+} ; oxide ion is O^{2-} . To balance the ionic charges, the formula of ferric oxide is Fe_2O_3 . [The systematic name is iron(III) oxide.]

(d) The anion is sulfide, S^{2-} , which requires that the cation be Co^{2+} . The name is cobalt(II) sulfide.

FOLLOW-UP PROBLEM 2.7 Give the systematic names for the formulas or the formulas for the names of the following compounds: (a) lead(IV) oxide; (b) Cu_2S ; (c) FeBr_2 ; (d) mercuric chloride.

Compounds Formed from Polyatomic Ions Ionic compounds in which one or both of the ions are polyatomic are very common. Table 2.5 gives the formulas and the names of some common polyatomic ions. Remember that *the polyatomic ion stays together as a charged unit*. The formula for potassium nitrate is KNO_3 ; each K^+ balances one NO_3^- . The formula for sodium carbonate is Na_2CO_3 ; two Na^+ balance one CO_3^{2-} . When two or more of the same polyatomic ion are

present in the formula unit, that ion appears in parentheses with the subscript written outside. For example, calcium nitrate, which contains one Ca^{2+} and two NO_3^- ions, has the formula $\text{Ca}(\text{NO}_3)_2$. Parentheses and a subscript are *not* used unless more than one of the polyatomic ions is present; thus, sodium nitrate is NaNO_3 , not $\text{Na}(\text{NO}_3)$.

Families of Oxoanions As Table 2.5 shows, most polyatomic ions are **oxoanions**, those in which an element, usually a nonmetal, is bonded to one or more oxygen atoms. There are several families of two or four oxoanions that differ only in the number of oxygen atoms. A simple naming convention is used with these ions.

With two oxoanions in the family:

- The ion with *more* O atoms takes the nonmetal root and the suffix *-ate*.
- The ion with *fewer* O atoms takes the nonmetal root and the suffix *-ite*.

For example, SO_4^{2-} is the sulf*ate* ion; SO_3^{2-} is the sulf*ite* ion; similarly, NO_3^- is nitr*ate*, and NO_2^- is nitr*ite*.

With four oxoanions in the family (usually a halogen bonded to O), as Figure 2.17 shows:

- The ion with *most* O atoms has the prefix *per-*, the nonmetal root, and the suffix *-ate*.
- The ion with *one fewer* O atom has just the root and the suffix *-ate*.
- The ion with *two fewer* O atoms has just the root and the suffix *-ite*.
- The ion with *least (three fewer)* O atoms has the prefix *hypo-*, the root, and the suffix *-ite*.

For example, for the four chlorine oxoanions,

ClO_4^- is *perchlorate*, ClO_3^- is chlor*ate*, ClO_2^- is chlor*ite*, ClO^- is *hypochlorite*

Hydrated Ionic Compounds Ionic compounds called **hydrates** have a specific number of water molecules associated with each formula unit. In their formulas, this number is shown after a centered dot. It is indicated in the systematic name by a Greek numerical prefix before the word *hydrate*. Table 2.6 shows these prefixes. For example, Epsom salt has the formula $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ and the name magnesium sulfate *heptahydrate*. Similarly, the mineral gypsum has the formula $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ and the name calcium sulfate *dihydrate*. The water molecules, referred to as “waters of hydration,” are part of the hydrate’s structure. Heating can remove some or all of them, leading to a different substance. For example, when heated strongly, blue copper(II) sulfate pentahydrate ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$) is converted to white copper(II) sulfate (CuSO_4).

	Prefix	Root	Suffix
No. of O atoms ↑	per	root	ate
		root	ate
		root	ite
	hypo	root	ite

Figure 2.17 Naming oxoanions. Prefixes and suffixes indicate the number of O atoms in the anion.

Table 2.6 Numerical Prefixes for Hydrates and Binary Covalent Compounds

Number	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

SAMPLE PROBLEM 2.8 Determining Names and Formulas of Ionic Compounds Containing Polyatomic Ions

Problem Give the systematic names for the formulas or the formulas for the names of the following compounds:

(a) $\text{Fe}(\text{ClO}_4)_2$ (b) Sodium sulfite (c) $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$

Solution (a) ClO_4^- is perchlorate, which has a 1− charge, so the cation must be Fe^{2+} . The name is **iron(II) perchlorate**. (The common name is ferrous perchlorate.)

(b) Sodium is Na^+ ; sulfite is SO_3^{2-} . Therefore, two Na^+ ions balance one SO_3^{2-} ion. The formula is **Na_2SO_3** . (c) Ba^{2+} is barium; OH^- is hydroxide. There are eight (*octa-*) water molecules in each formula unit. The name is **barium hydroxide octahydrate**.

FOLLOW-UP PROBLEM 2.8 Give the systematic names for the formulas or the formulas for the names of the following compounds:

(a) Cupric nitrate trihydrate (b) Zinc hydroxide (c) LiCN

SAMPLE PROBLEM 2.9 Recognizing Incorrect Names and Formulas of Ionic Compounds

Problem Something is wrong with the second part of each statement. Provide the correct name or formula.

- (a) $\text{Ba}(\text{C}_2\text{H}_3\text{O}_2)_2$ is called barium diacetate.
 (b) Sodium sulfide has the formula $(\text{Na})_2\text{SO}_3$.
 (c) Iron(II) sulfate has the formula $\text{Fe}_2(\text{SO}_4)_3$.
 (d) Cesium carbonate has the formula $\text{Cs}_2(\text{CO}_3)$.

Solution (a) The charge of the Ba^{2+} ion *must* be balanced by *two* $\text{C}_2\text{H}_3\text{O}_2^-$ ions, so the prefix *di-* is unnecessary. For ionic compounds, we do not indicate the number of ions with numerical prefixes. The correct name is **barium acetate**.

(b) Two mistakes occur here. The sodium ion is monatomic, so it does *not* require parentheses. The sulfide ion is S^{2-} , *not* SO_3^{2-} (called “sulfite”). The correct formula is **Na_2S** .

(c) The Roman numeral refers to the charge of the ion, *not* the number of ions in the formula. Fe^{2+} is the cation, so it requires one SO_4^{2-} to balance its charge. The correct formula is **FeSO_4** .

(d) Parentheses are *not* required when only one polyatomic ion of a kind is present. The correct formula is **Cs_2CO_3** .

FOLLOW-UP PROBLEM 2.9 State why the second part of each statement is incorrect, and correct it:

- (a) Ammonium phosphate is $(\text{NH}_3)_4\text{PO}_4$. (b) Aluminum hydroxide is AlOH_3 .
 (c) $\text{Mg}(\text{HCO}_3)_2$ is manganese(II) carbonate. (d) $\text{Cr}(\text{NO}_3)_3$ is chromic(III) nitride.
 (e) $\text{Ca}(\text{NO}_2)_2$ is cadmium nitrate.

Acid Names from Anion Names Acids are an important group of hydrogen-containing compounds that have been used in chemical reactions for centuries. In the laboratory, acids are typically used in water solution. When naming them and writing their formulas, we consider them as anions connected to the number of hydrogen ions (H^+) needed for charge neutrality. The two common types of acids are binary acids and oxoacids:

1. *Binary acid* solutions form when certain gaseous compounds dissolve in water. For example, when gaseous hydrogen chloride (HCl) dissolves in water, it forms a solution whose name consists of the following parts:

Prefix *hydro-* + nonmetal *root* + suffix *-ic* + separate word *acid*
 hydro + chlor + ic + acid

or *hydrochloric acid*. This naming pattern holds for many compounds in which hydrogen combines with an anion that has an *-ide* suffix.

2. *Oxoacid* names are similar to those of the oxoanions, except for two suffix changes:

- *-ate* in the anion becomes *-ic* in the acid
- *-ite* in the anion becomes *-ous* in the acid

The oxoanion prefixes *hypo-* and *per-* are kept. Thus,

BrO_4^- is *perbromate*, and HBrO_4 is *perbromic acid*
 IO_2^- is *iodite*, and HIO_2 is *iodous acid*

SAMPLE PROBLEM 2.10 Determining Names and Formulas of Anions and Acids

Problem Name the following anions and give the names and formulas of the acids derived from them: (a) Br^- ; (b) IO_3^- ; (c) CN^- ; (d) SO_4^{2-} ; (e) NO_2^- .

Solution (a) The anion is **bromide**; the acid is **hydrobromic acid, HBr** .

(b) The anion is **iodate**; the acid is **iodic acid, HIO_3** .

(c) The anion is **cyanide**; the acid is **hydrocyanic acid, HCN** .

(d) The anion is sulfate; the acid is sulfuric acid, H_2SO_4 . (In this case, the suffix is added to the element name *sulfur*, not to the root, *sulf-*.)

(e) The anion is nitrite; the acid is nitrous acid, HNO_2 .

Comment We added *two* H^+ ions to the sulfate ion to obtain sulfuric acid because it has a $2-$ charge.

FOLLOW-UP PROBLEM 2.10 Write the formula for the name or name for the formula of each acid: (a) chloric acid; (b) HF ; (c) acetic acid; (d) sulfurous acid; (e) HBrO .

Names and Formulas of Binary Covalent Compounds

Binary covalent compounds are formed by the combination of two elements, usually nonmetals. Several are so familiar, such as ammonia (NH_3), methane (CH_4), and water (H_2O), that we use their common names, but most are named in a systematic way:

1. The element with the lower group number in the periodic table is the first word in the name; the element with the higher group number is the second word. (*Exception:* When the compound contains oxygen and any of the halogens chlorine, bromine, and iodine, the halogen is named first.)
2. If both elements are in the same group, the one with the higher period number is named first.
3. The second element is named with its root and the suffix *-ide*.
4. Covalent compounds have Greek numerical prefixes (see Table 2.6) to indicate the number of atoms of each element in the compound. The first word has a prefix *only* when more than one atom of the element is present; the second word *usually* has a numerical prefix.

SAMPLE PROBLEM 2.11 Determining Names and Formulas of Binary Covalent Compounds

Problem (a) What is the formula of carbon disulfide? (b) What is the name of PCl_5 ? (c) Give the name and formula of the compound whose molecules each consist of two N atoms and four O atoms.

Solution (a) The prefix *di-* means “two.” The formula is CS_2 .

(b) P is the symbol for phosphorus; there are five chlorine atoms, which is indicated by the prefix *penta-*. The name is phosphorus pentachloride.

(c) Nitrogen (N) comes first in the name (lower group number). The compound is dinitrogen tetroxide, N_2O_4 .

FOLLOW-UP PROBLEM 2.11 Give the name or formula for (a) SO_3 ; (b) SiO_2 ; (c) dinitrogen monoxide; (d) selenium hexafluoride.

SAMPLE PROBLEM 2.12 Recognizing Incorrect Names and Formulas of Binary Covalent Compounds

Problem Explain what is wrong with the name or formula in the second part of each statement and correct it: (a) SF_4 is monosulfur pentafluoride. (b) Dichlorine heptaoxide is Cl_2O_6 . (c) N_2O_3 is dinitrotrioxide.

Solution (a) There are two mistakes. *Mono-* is not needed if there is only one atom of the first element, and the prefix for four is *tetra-*, not *penta-*. The correct name is sulfur tetrafluoride.

(b) The prefix *hepta-* indicates seven, not six. The correct formula is Cl_2O_7 .

(c) The full name of the first element is needed, and a space separates the two element names. The correct name is dinitrogen trioxide.

FOLLOW-UP PROBLEM 2.12 Explain what is wrong with the second part of each statement and correct it: (a) S_2Cl_2 is disulfurous dichloride. (b) Nitrogen monoxide is N_2O . (c) BrCl_3 is trichlorine bromide.

Naming Alkanes

Organic compounds typically have complex structural formulas that consist of chains, branches, and/or rings of carbon atoms bonded to hydrogen atoms and, often, to atoms of oxygen, nitrogen, and a few other elements. At this point, we'll see how the simplest organic compounds are named. Much more on the rules of organic nomenclature appears in Chapter 15.

Hydrocarbons, the simplest type of organic compound, contain *only* carbon and hydrogen. *Alkanes* are the simplest type of hydrocarbon; many function as important fuels, such as methane, propane, butane, and the mixture of alkanes in gasoline. The simplest alkanes to name are the *straight-chain alkanes* because the carbon chains have no branches. Alkanes are named with a *root*, based on the number of C atoms in the chain, followed by the suffix *-ane*. Table 2.7 gives the names, molecular formulas, and space-filling models (discussed shortly) of the first 10 straight-chain alkanes. Note that the roots of the four smallest ones are new, but those for the larger ones are the same as the Greek prefixes in Table 2.6.

Molecular Masses from Chemical Formulas

In Section 2.5, we calculated the atomic mass of an element. Using the periodic table and the formula of a compound to see the number of atoms of each element present, we calculate the **molecular mass** (also called *molecular weight*) of a formula unit of the compound as the sum of the atomic masses:

$$\text{Molecular mass} = \text{sum of atomic masses} \quad (2.3)$$

The molecular mass of a water molecule (using atomic masses to four significant figures from the periodic table) is

$$\begin{aligned} \text{Molecular mass of H}_2\text{O} &= (2 \times \text{atomic mass of H}) + (1 \times \text{atomic mass of O}) \\ &= (2 \times 1.008 \text{ amu}) + 16.00 \text{ amu} = 18.02 \text{ amu} \end{aligned}$$

Ionic compounds are treated the same, but because they do not consist of molecules, we use the term **formula mass** for an ionic compound. To calculate its formula mass, *the number of atoms of each element inside the parentheses is multiplied by the subscript outside the parentheses*. For barium nitrate, $\text{Ba}(\text{NO}_3)_2$,

Formula mass of $\text{Ba}(\text{NO}_3)_2$

$$\begin{aligned} &= (1 \times \text{atomic mass of Ba}) + (2 \times \text{atomic mass of N}) + (6 \times \text{atomic mass of O}) \\ &= 137.3 \text{ amu} + (2 \times 14.01 \text{ amu}) + (6 \times 16.00 \text{ amu}) = 261.3 \text{ amu} \end{aligned}$$

Note that atomic, not ionic, masses are used. Although masses of ions differ from those of their atoms by the masses of the electrons, electron loss equals electron gain in the compound, so electron mass is balanced.

SAMPLE PROBLEM 2.13 Calculating the Molecular Mass of a Compound











Problem Using data in the periodic table, calculate the molecular (or formula) mass of: (a) Tetraphosphorus trisulfide (b) Ammonium nitrate

Plan We first write the formula, then multiply the number of atoms (or ions) of each element by its atomic mass, and find the sum.

Solution (a) The formula is P_4S_3 .

$$\begin{aligned} \text{Molecular mass} &= (4 \times \text{atomic mass of P}) + (3 \times \text{atomic mass of S}) \\ &= (4 \times 30.97 \text{ amu}) + (3 \times 32.07 \text{ amu}) = 220.09 \text{ amu} \end{aligned}$$

Table 2.7 The First 10 Straight-Chain Alkanes

Name (Formula)	Model
Methane (CH_4)	
Ethane (C_2H_6)	
Propane (C_3H_8)	
Butane (C_4H_{10})	
Pentane (C_5H_{12})	
Hexane (C_6H_{14})	
Heptane (C_7H_{16})	
Octane (C_8H_{18})	
Nonane (C_9H_{20})	
Decane ($\text{C}_{10}\text{H}_{22}$)	

(b) The formula is NH_4NO_3 . We count the total number of N atoms even though they belong to different ions:

Formula mass

$$= (2 \times \text{atomic mass of N}) + (4 \times \text{atomic mass of H}) + (3 \times \text{atomic mass of O})$$

$$= (2 \times 14.01 \text{ amu}) + (4 \times 1.008 \text{ amu}) + (3 \times 16.00 \text{ amu}) = 80.05 \text{ amu}$$

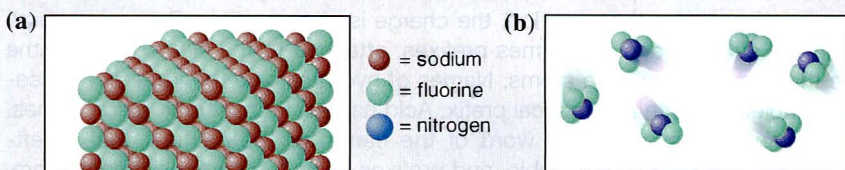
Check You can often find large errors by rounding atomic masses to the nearest 5 and adding: (a) $(4 \times 30) + (3 \times 30) = 210 \approx 220.09$. The sum has two decimal places because the atomic masses have two. (b) $(2 \times 15) + 4 + (3 \times 15) = 79 \approx 80.05$.

FOLLOW-UP PROBLEM 2.13 What is the formula and molecular (or formula) mass of each of the following compounds: (a) hydrogen peroxide; (b) cesium chloride; (c) sulfuric acid; (d) potassium sulfate?

In the next sample problem, we use molecular depictions to find the formula, name, and mass.

SAMPLE PROBLEM 2.14 Determining Formulas and Names from Molecular Depictions

Problem Each box contains a representation of a binary compound. Determine its formula, name, and molecular (formula) mass.



Plan Each of the compounds contains only two elements, so to find the formula, we find the simplest whole-number ratio of one atom to the other. Then we determine the name (see Sample Problems 2.5, 2.6, and 2.11) and the mass (see Sample Problem 2.13).

Solution (a) There is one brown (sodium) for each green (fluorine), so the formula is NaF . A metal and nonmetal form an ionic compound, in which the metal is named first: sodium fluoride.

$$\begin{aligned} \text{Formula mass} &= (1 \times \text{atomic mass of Na}) + (1 \times \text{atomic mass of F}) \\ &= 22.99 \text{ amu} + 19.00 \text{ amu} = 41.99 \text{ amu} \end{aligned}$$

(b) There are three green (fluorine) for each blue (nitrogen), so the formula is NF_3 . Two nonmetals form a covalent compound. Nitrogen has a lower group number, so it is named first: nitrogen trifluoride.

$$\begin{aligned} \text{Molecular mass} &= (1 \times \text{atomic mass of N}) + (3 \times \text{atomic mass of F}) \\ &= 14.01 \text{ amu} + (3 \times 19.00 \text{ amu}) = 71.01 \text{ amu} \end{aligned}$$

Check (a) For binary ionic compounds, we predict ionic charges from the periodic table (see Figure 2.10). Na forms a $1+$ ion, and F forms a $1-$ ion, so the charges balance with one Na^+ per F^- . Also, ionic compounds are solids, consistent with the picture. (b) Covalent compounds often occur as individual molecules, as in the picture. Rounding in (a) gives $25 + 20 = 45$; in (b), we get $15 + (3 \times 20) = 75$, so there are no large errors.

FOLLOW-UP PROBLEM 2.14 Each box contains a representation of a binary compound. Determine its name, formula, and molecular (formula) mass.

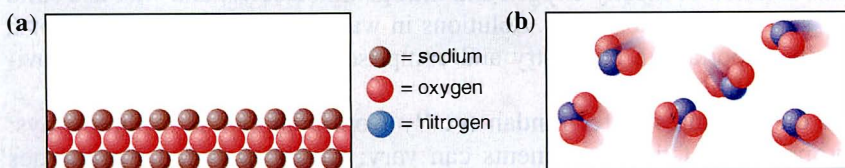
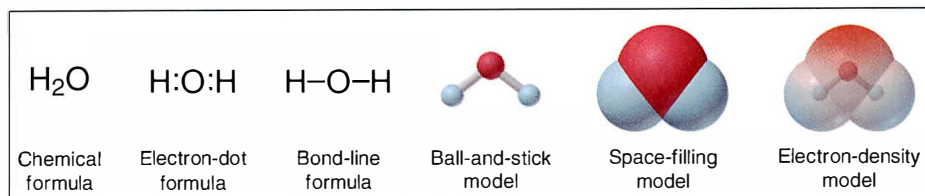


Figure 2.18 Representations of a water molecule.



Picturing Molecules

Molecules are depicted in a variety of useful ways, as shown in Figure 2.18 for the water molecule. A *chemical formula* shows only the relative number of atoms. *Electron-dot* and *bond-line formulas* show a bond between atoms as either a pair of dots or a line. A *ball-and-stick model* shows atoms as spheres and bonds as sticks, with accurate angles and relative sizes, but distances are exaggerated. A *space-filling model* is an accurately scaled-up version of a molecule, but it does not show bonds. An *electron-density model* shows the ball-and-stick model within the space-filling shape and colors the regions of high (*red*) and low (*blue*) electron charge.

SECTION SUMMARY

Chemical formulas describe the simplest atom ratio (empirical formula), actual atom number (molecular formula), and atom arrangement (structural formula) of one unit of a compound. An ionic compound is named with cation first and anion second. For metals that can form more than one ion, the charge is shown with a Roman numeral. Oxoanions have suffixes, and sometimes prefixes, attached to the element root name to indicate the number of oxygen atoms. Names of hydrates give the number of associated water molecules with a numerical prefix. Acid names are based on anion names. Covalent compounds have the first word of the name for the element that is left-most or lower down in the periodic table, and prefixes show the number of each atom. The molecular (or formula) mass of a compound is the sum of the atomic masses in the formula. Molecules are depicted by various types of formulas and models.

2.9 CLASSIFICATION OF MIXTURES

Although chemists pay a great deal of attention to pure substances, this form of matter almost never occurs around us. In the natural world, *matter usually occurs as mixtures*, such as air, seawater, soil, and organisms.

There are two broad classes of mixtures. A **heterogeneous mixture** has one or more visible boundaries between the components. Thus, its composition is *not* uniform; it varies from one region to another. Many rocks are heterogeneous, showing individual grains and flecks of different minerals. In some cases, as in milk and blood, the boundaries can be seen only with a microscope. A **homogeneous mixture** has no visible boundaries because the components are mixed as individual atoms, ions, and molecules. Thus, its composition *is* uniform. A mixture of sugar dissolved in water is homogeneous, for example, because the sugar molecules and water molecules are uniformly intermingled on the molecular level. We have no way to tell visually whether an object is a substance (element or compound) or a homogeneous mixture.

A homogeneous mixture is also called a **solution**. Although we usually think of solutions as liquid, they can exist in all three physical states. For example, air is a gaseous solution of mostly oxygen and nitrogen molecules, and wax is a solid solution of several fatty substances. Solutions in water, called **aqueous solutions**, are especially important in chemistry and comprise a major portion of the environment and of all organisms.

Recall that mixtures differ fundamentally from compounds in three ways: (1) the proportions of the components can vary; (2) the individual properties of the components are observable; and (3) the components can be separated by physical means. In some cases, as in a mixture of iron and sulfur (Figure 2.19),

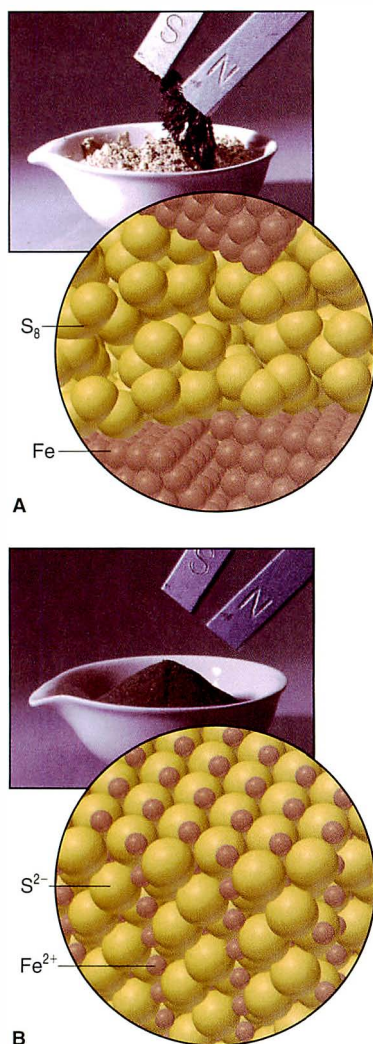


Figure 2.19 The distinction between mixtures and compounds. **A**, A mixture of iron and sulfur can be separated with a magnet because only the iron is magnetic. The blow-up shows separate regions of the two elements. **B**, After strong heating, the compound iron(II) sulfide forms, which is no longer magnetic. The blow-up shows the structure of the compound.

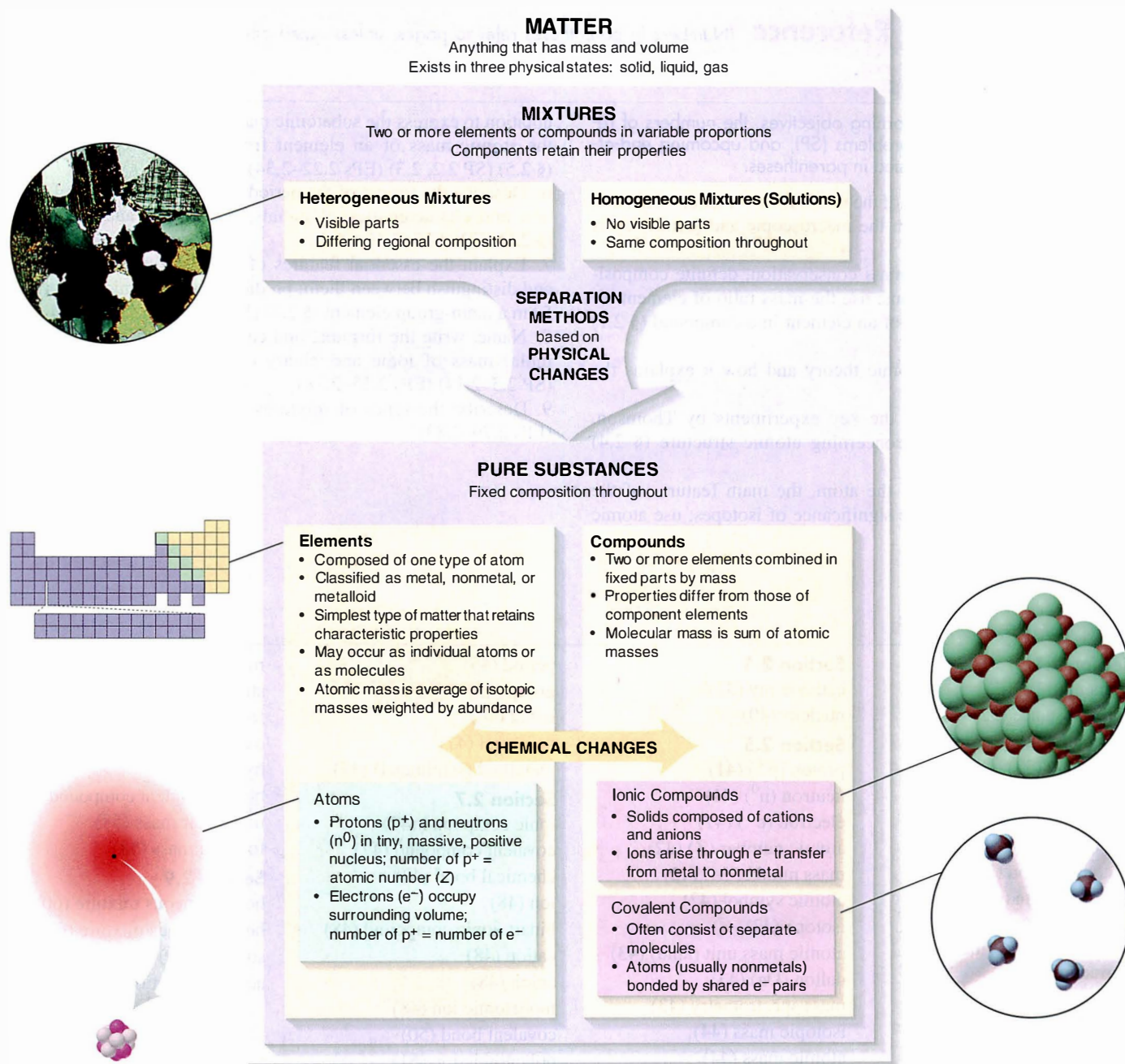


Figure 2.20 The classification of matter from a chemical point of view. Mixtures are separated by physical changes into elements and

compounds. Chemical changes are required to convert elements into compounds, and vice versa.

if we apply enough energy to the components, they react with each other and form a compound, after which their individual properties are no longer observable. The characteristics of mixtures and pure substances that we covered in this chapter are summarized in Figure 2.20.

SECTION SUMMARY

Heterogeneous mixtures have visible boundaries between the components. Homogeneous mixtures have no visible boundaries because mixing occurs at the molecular level. A solution is a homogeneous mixture and can occur in any physical state. Mixtures (not compounds) can have variable proportions, can be separated physically, and retain their components' properties.